

# QUESTION

Suppose the weak acid  $\text{HNO}_2$  ( $K_a = 4.0 \times 10^{-4}$ ) was added to a solution of  $\text{NaNO}_2$ . If the concentration of acid were 0.10 M and the salt concentration was 0.060 M, what would be the  $[\text{H}^+]$ ?

$$\text{Ans} = 6.7 \times 10^{-4} \text{ M}$$

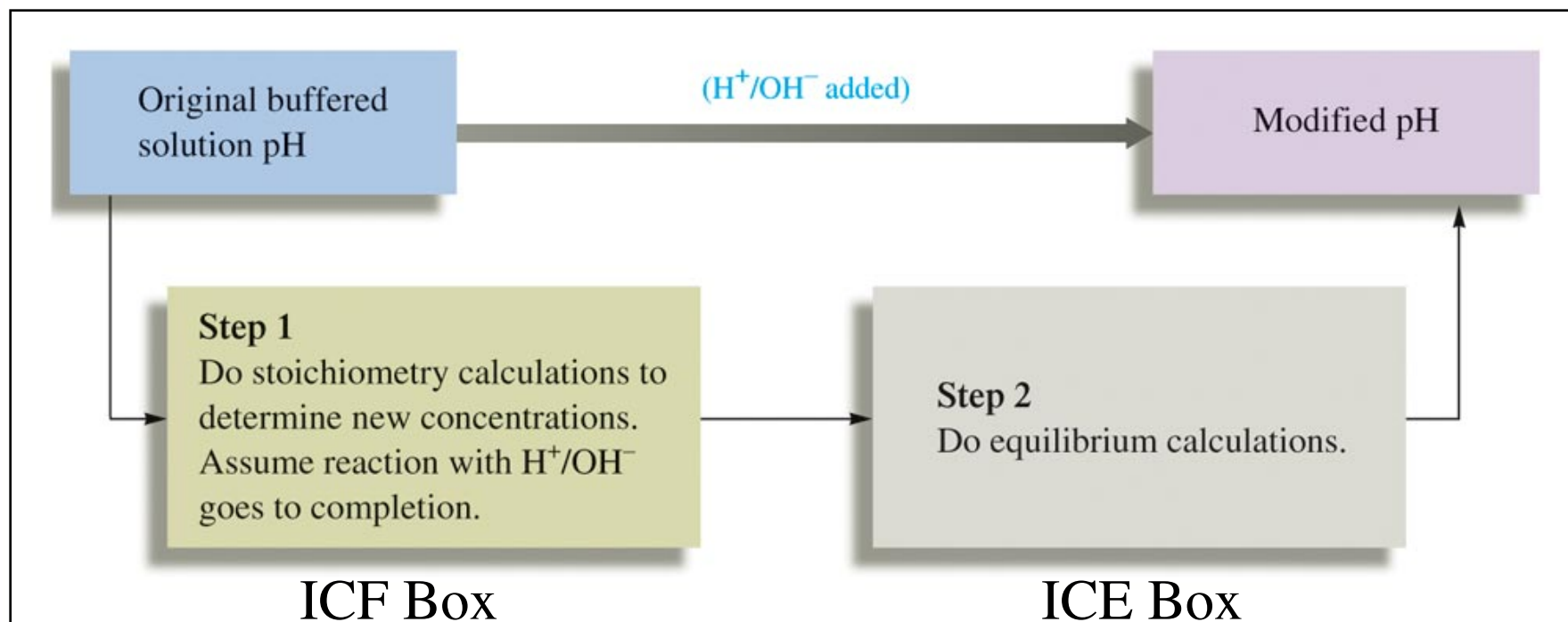
Compare this to  $[\text{H}^+]$  of just acid with no conjugate base.

# QUESTION

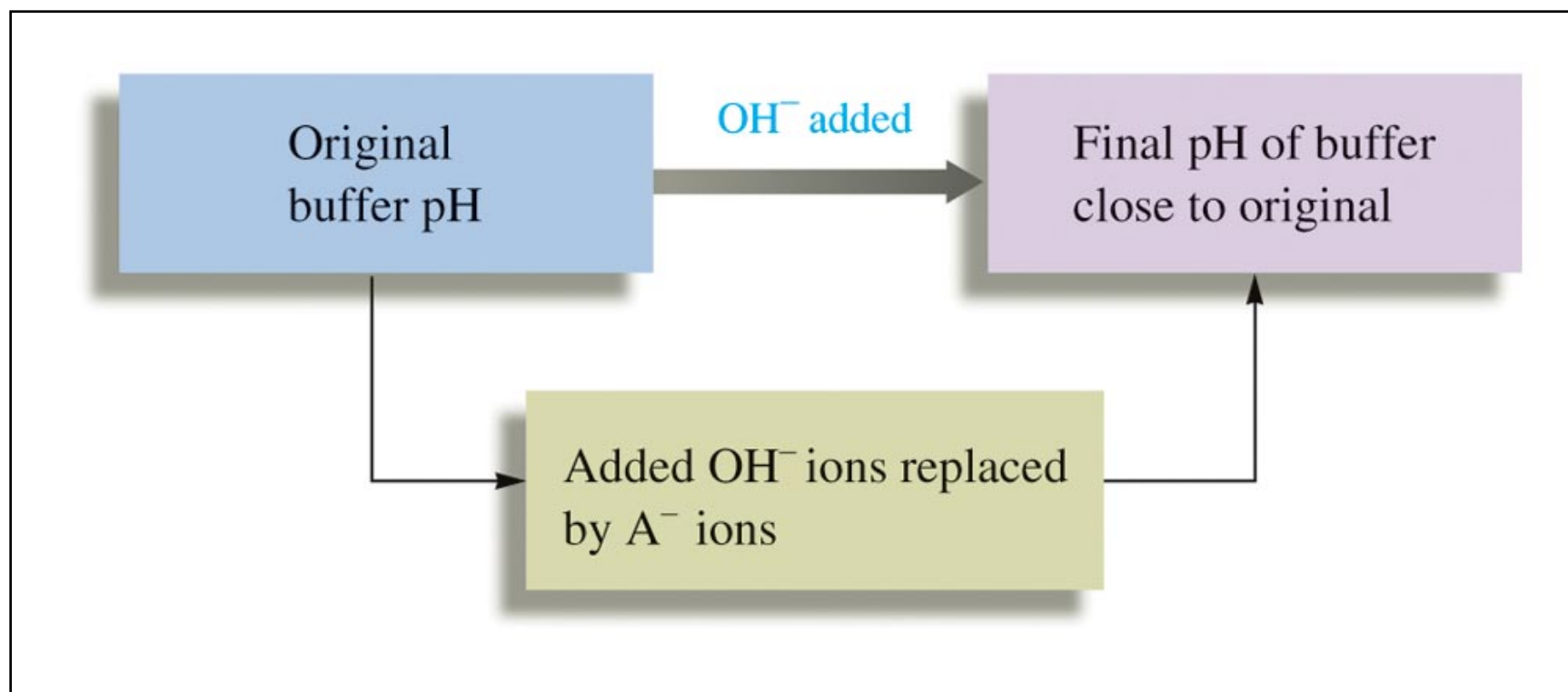
A 0.50 L buffer solution containing 0.42 M  $\text{NH}_4\text{Cl}$  and 0.75 M  $\text{NH}_3$  ( $K_b$  of  $\text{NH}_3 = 1.8 \times 10^{-5}$ ) has a pH of 9.51 at 25°C. The solution receives 0.010 moles of HCl from an outside source. Assuming no significant change in volume of the solution, what would be the pH of the solution after the addition of the HCl at 25°C?

Ans = 9.47

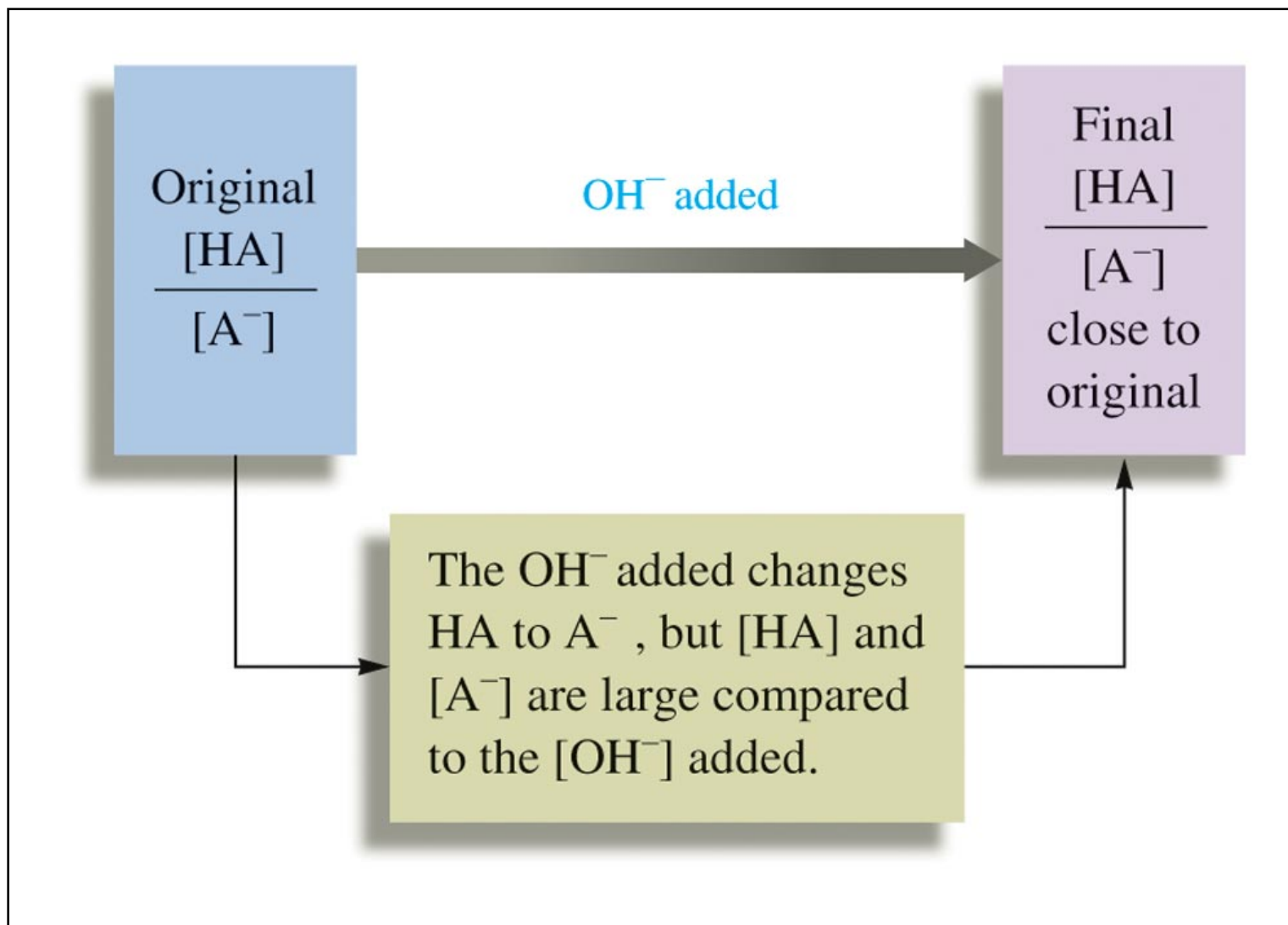
# Strong Acid or Base Being Added to a Buffered Solution



# How Buffering Works



# Buffering Schematic



# QUESTION

Calculate the pH change that occurs when 0.010 mol of solid NaOH is added to 1.0 L of the 0.50 M acetic acid and 0.50 M sodium acetate? ( $K_a = 1.8 \times 10^{-5}$ )

Ans = +0.02

HH

# QUESTION

Calculate the change in pH that occurs when 0.010 mol of HCl is added to 1.0L of each of the following solutions:

Solution A: 5.00 M HOAc and 5.00 M NaOAc

Solution B: 0.050 M HOAc and 0.050 NaOAc

Assume no volume changes,  $K_a = 1.8 \times 10^{-5}$

Ans A = 0

B = -0.18

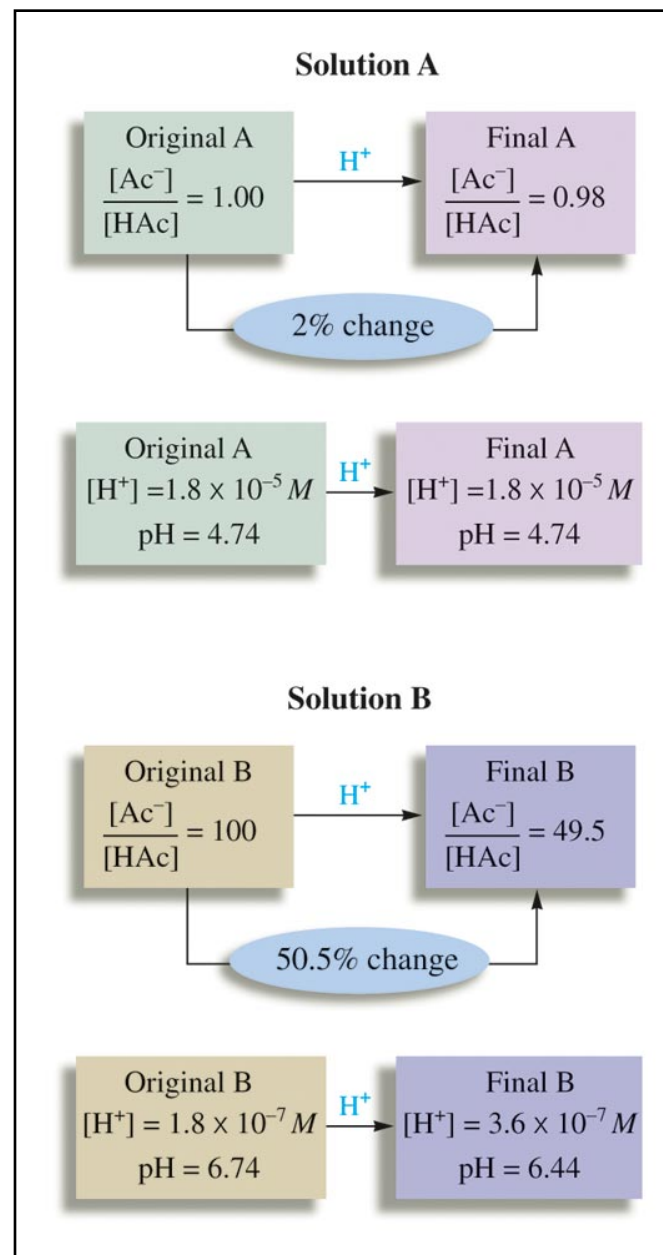
HH

# Table 15.1 Change in $[\text{C}_2\text{H}_3\text{O}_2^-]/[\text{HC}_2\text{H}_3\text{O}_2^-]$ for Two solutions When 0.01 mol $\text{H}^+$ is Added to 1.0 L of each

**TABLE 15.1 Change in  $[\text{C}_2\text{H}_3\text{O}_2^-]/[\text{HC}_2\text{H}_3\text{O}_2^-]$  for Two Solutions When 0.01 mol  $\text{H}^+$  Is Added to 1.0 L of Each**

Solution	$\left(\frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}\right)_{\text{orig}}$	$\left(\frac{[\text{C}_2\text{H}_3\text{O}_2^-]}{[\text{HC}_2\text{H}_3\text{O}_2]}\right)_{\text{new}}$	Change	Percent Change
A	$\frac{1.00\text{ M}}{1.00\text{ M}} = 1.00$	$\frac{0.99\text{ M}}{1.01\text{ M}} = 0.98$	$1.00 \rightarrow 0.98$	2.00%
B	$\frac{1.00\text{ M}}{0.01\text{ M}} = 100$	$\frac{0.99\text{ M}}{0.02\text{ M}} = 49.5$	$100 \rightarrow 49.5$	50.5%

# Soution A; Solution B



# QUESTION

To culture a certain bacteria a microbiologist would like to buffer the media at a pH of 3.75. To maximize the efficiency of the system a 1:1 ratio of acid to salt will be used. Which of the following acids would make the best choice for the buffer?

1. Acetic acid;  $K_a = 1.8 \times 10^{-5}$
2. Propanoic acid;  $K_a = 1.3 \times 10^{-5}$
3. Formic acid;  $K_a = 1.8 \times 10^{-4}$
4. Nitrous acid;  $K_a = 4.0 \times 10^{-4}$

# QUESTION

What mass of NaBrO should be added to 1.0 L 0.050 M HBrO to form a buffer with a pH of 9.15? Assume no volume changes.

$$K_a = 2.5 \times 10^{-9}$$

$$\text{Ans} = 21.4 \text{ g}$$

HH

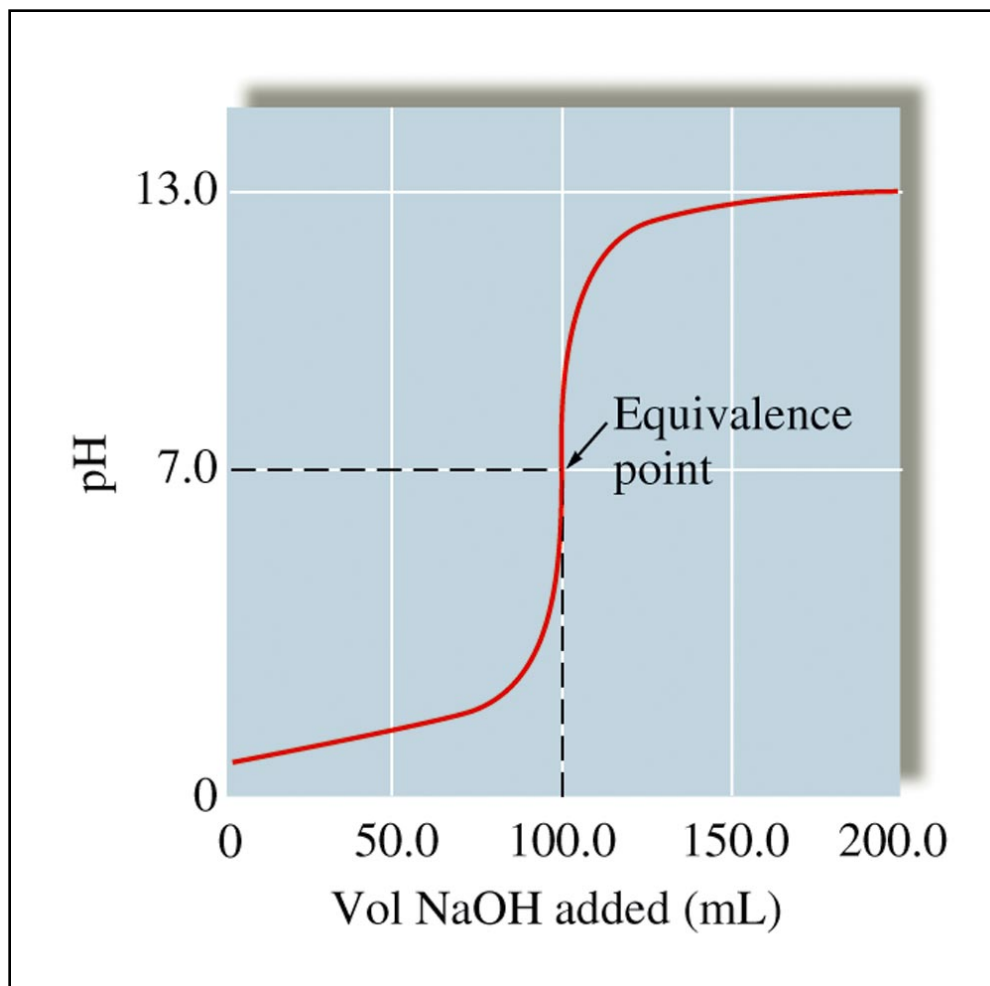
# QUESTION

What volume of 0.100 M NaOH is needed to neutralize 100 mL of 0.050 M  $\text{NaH}_2\text{PO}_4$  to get a buffer of pH 8.00? ( $K_a = 6.2 \times 10^{-8}$ )

Ans = 43 mL

HH

Figure 15.1 The pH Curve for the Titration of 50.0mL of 0.200 *M* HCl with 0.100 *M* NaOH



## Figure 15.2 The pH Curve for the Titration of 100.0 mL of 0.50 *M* NaOH with 1.0 *M* HCl

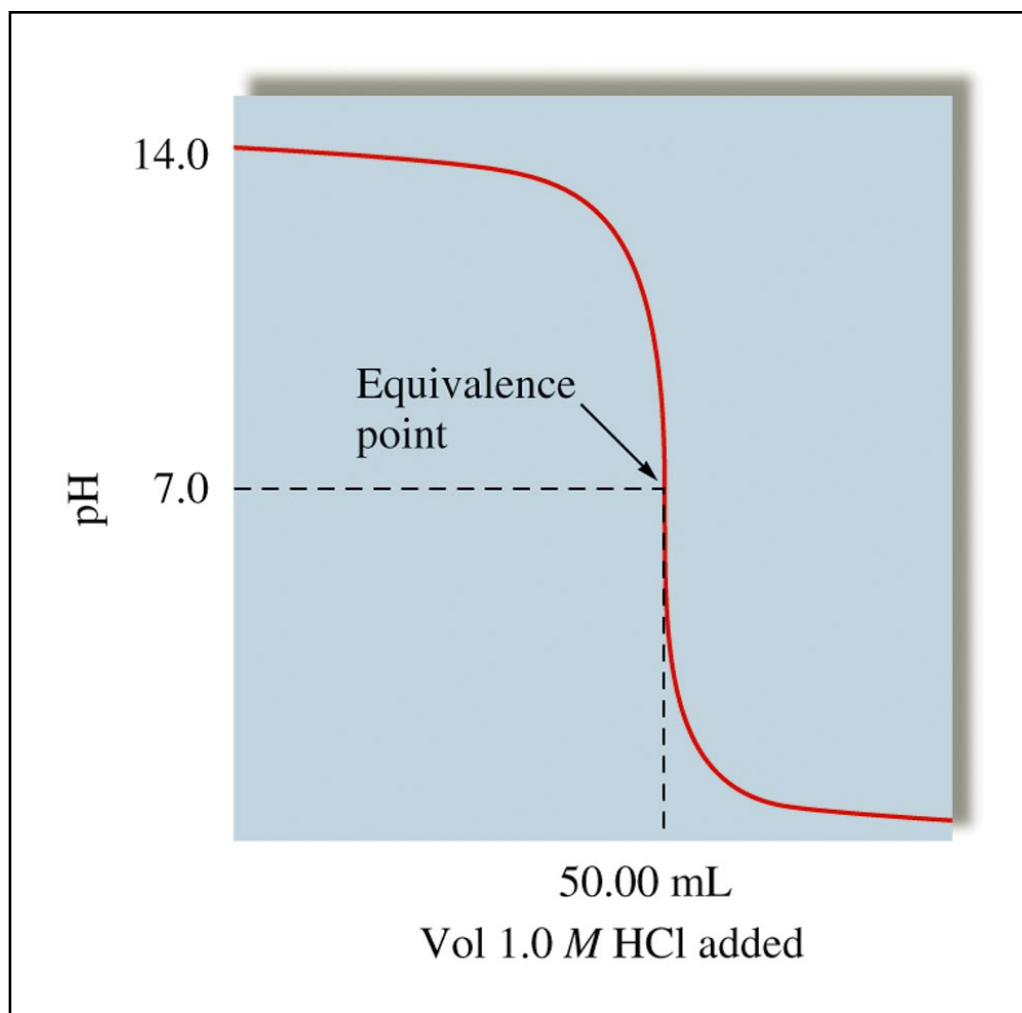
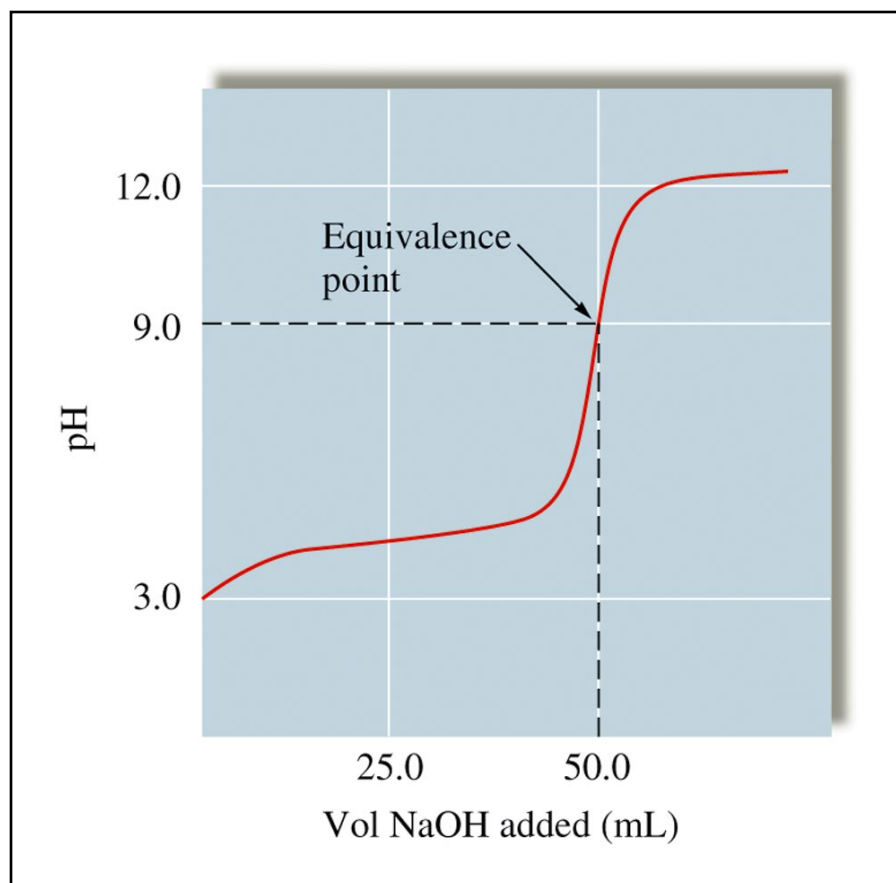
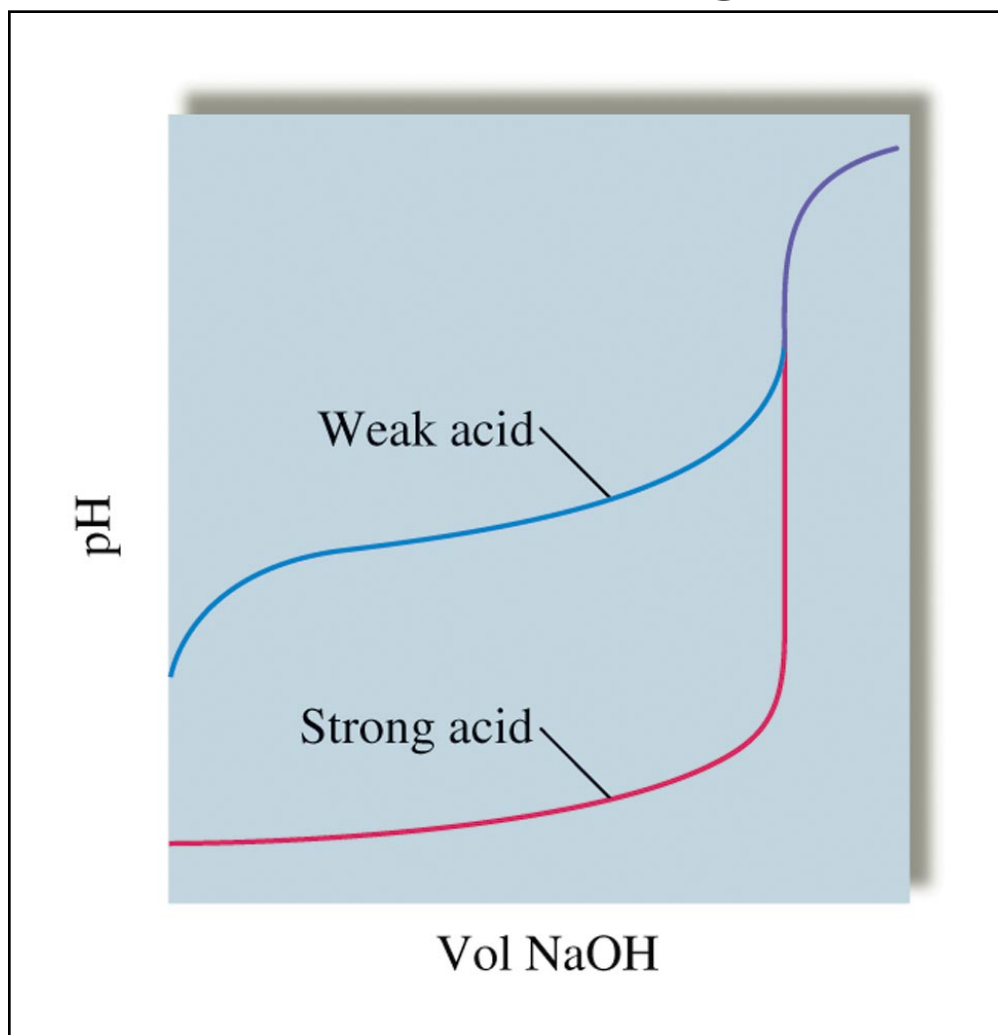


Figure 15.3 The pH Curve for the Titration of 50.0 mL of 0.100 M  $\text{HC}_2\text{H}_3\text{O}_2$  with 0.100 M  $\text{NaOH}$  Occurs at 50.0 mL of  $\text{NaOH}$  Added, Where the Amount of Added  $\text{OH}^-$  Exactly Equals the Original Amount of Acid



# Equivalence Points for the Titrations of Weak and Strong Acids



# QUESTION

If 50.0 mL of 0.100 M HCN ( $K_a = 6.2 \times 10^{-10}$ ) is titrated with 0.100 M NaOH, calculate the pH of solution:

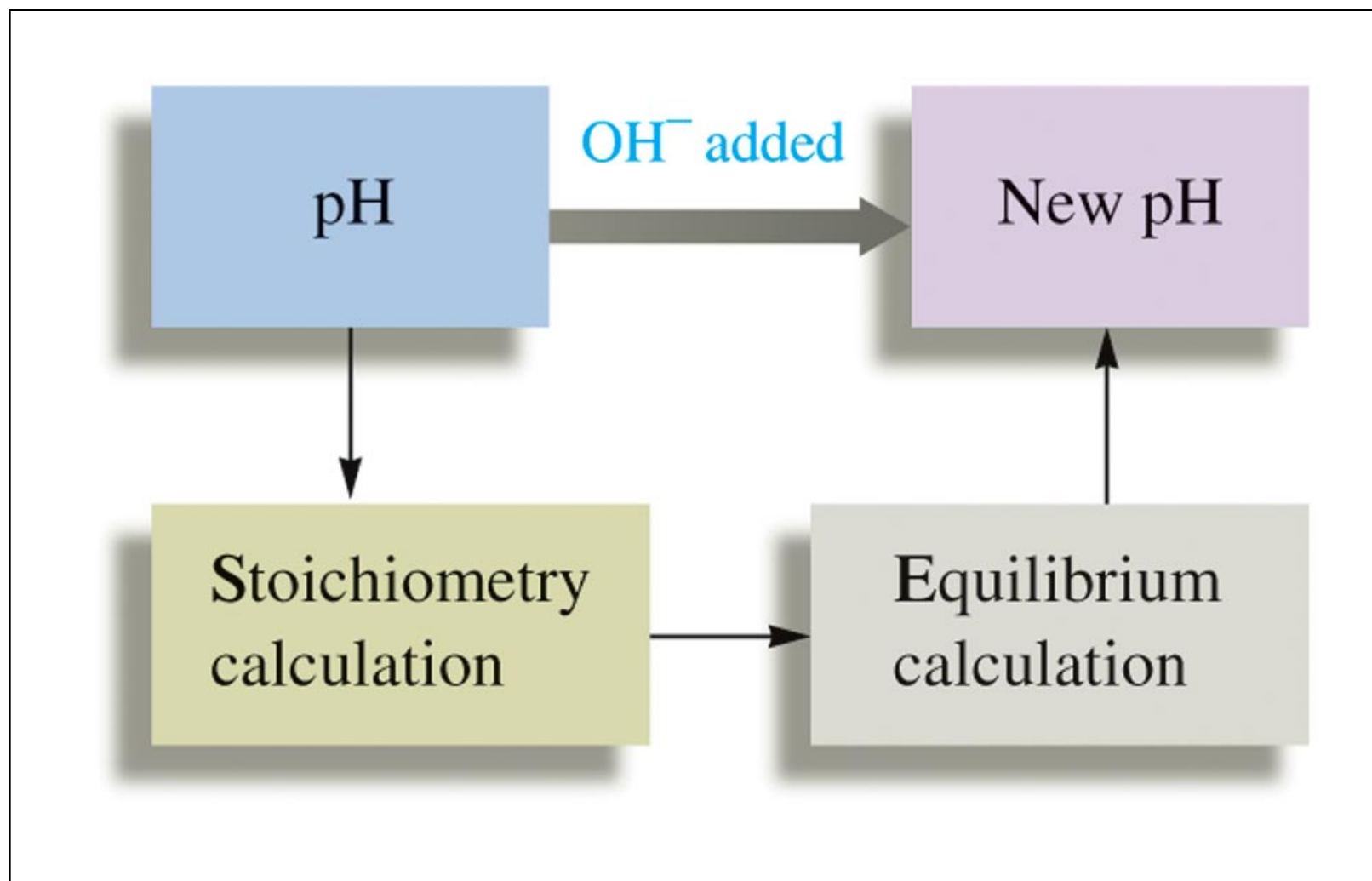
- After 8.00 mL of NaOH have been added.
- At the half-way point of the titration.
- At the equivalence point of the titration.

$$A = 8.49$$

$$B = 9.21$$

$$C = 10.96$$

## Treat the Stoichiometry and Equilibrium Problems Separately for Weak Acid/Strong Base



## Figure 15.4 Acid Strength Affects pH at Equivalence Point

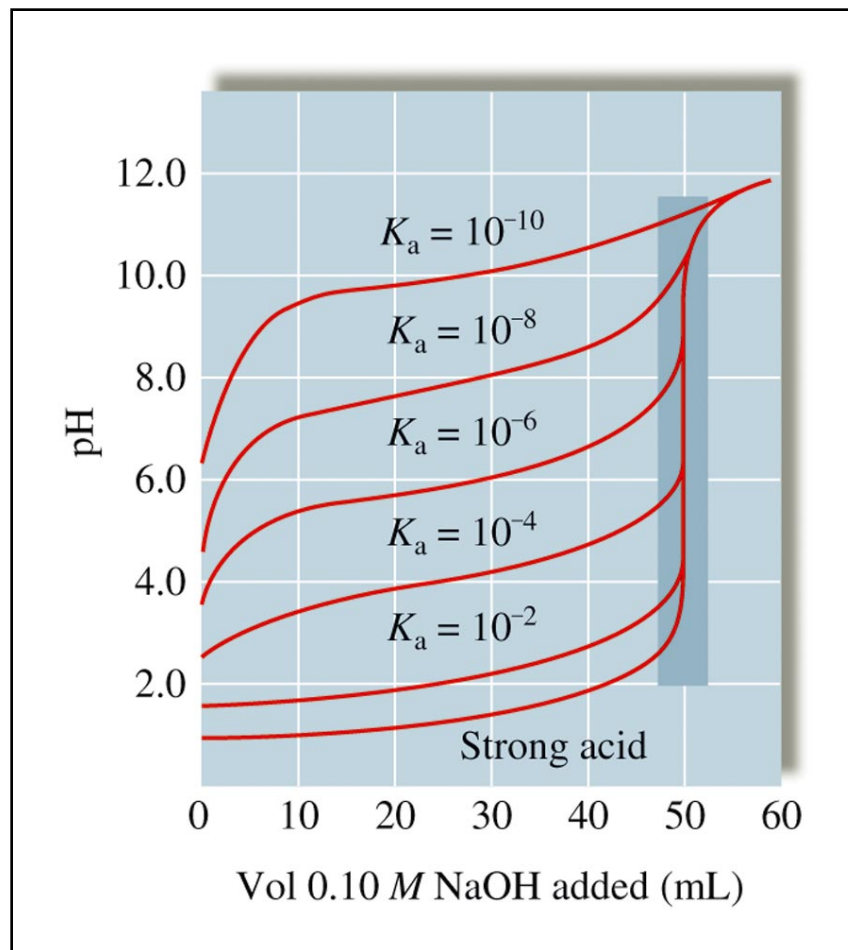
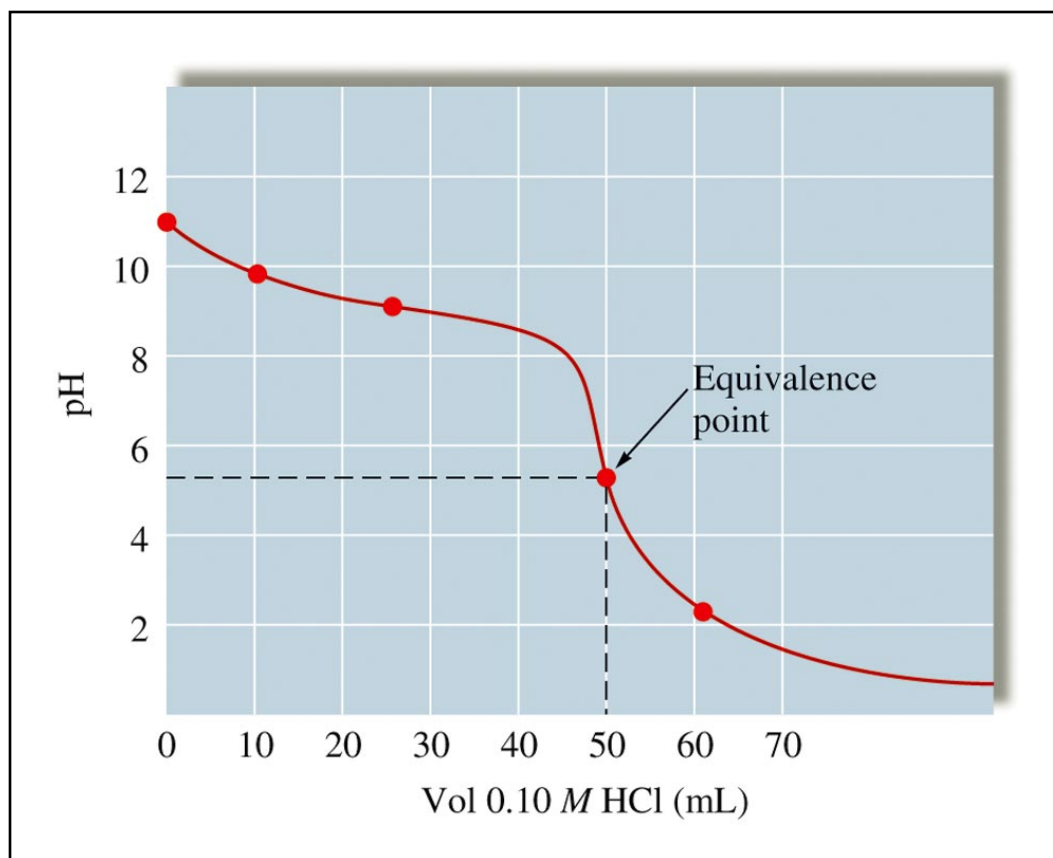
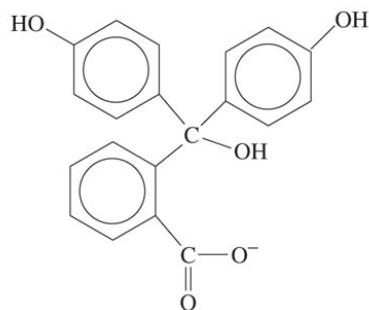
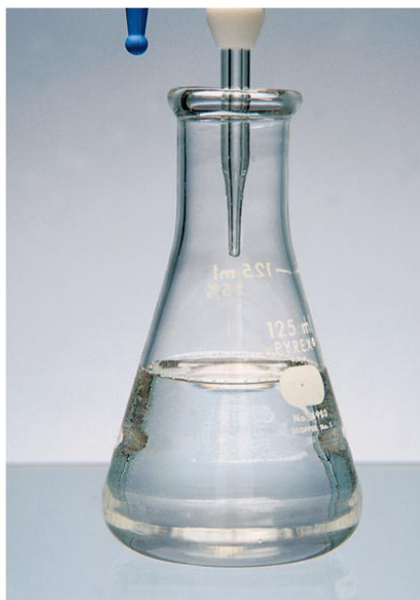


Figure 15.5 The pH Curve for the Titration of 100.0 mL of 0.050  $M$   $\text{NH}_3$  with 0.10  $M$   $\text{HCl}$ . Note the pH at the Equivalence Point is Less than 7, Since the Solution Contains the Weak Acid  $\text{NH}_4^+$

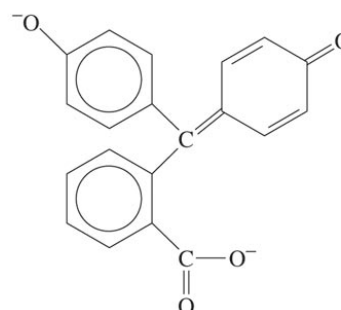
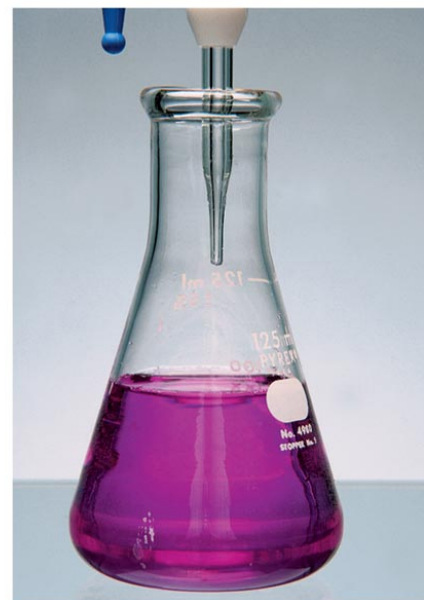


# Figure 15.6 The Acid and Base Forms of the Indicator Phenolphthalein

**END POINT  
IS NOT  
EQUIVALENCE  
POINT**

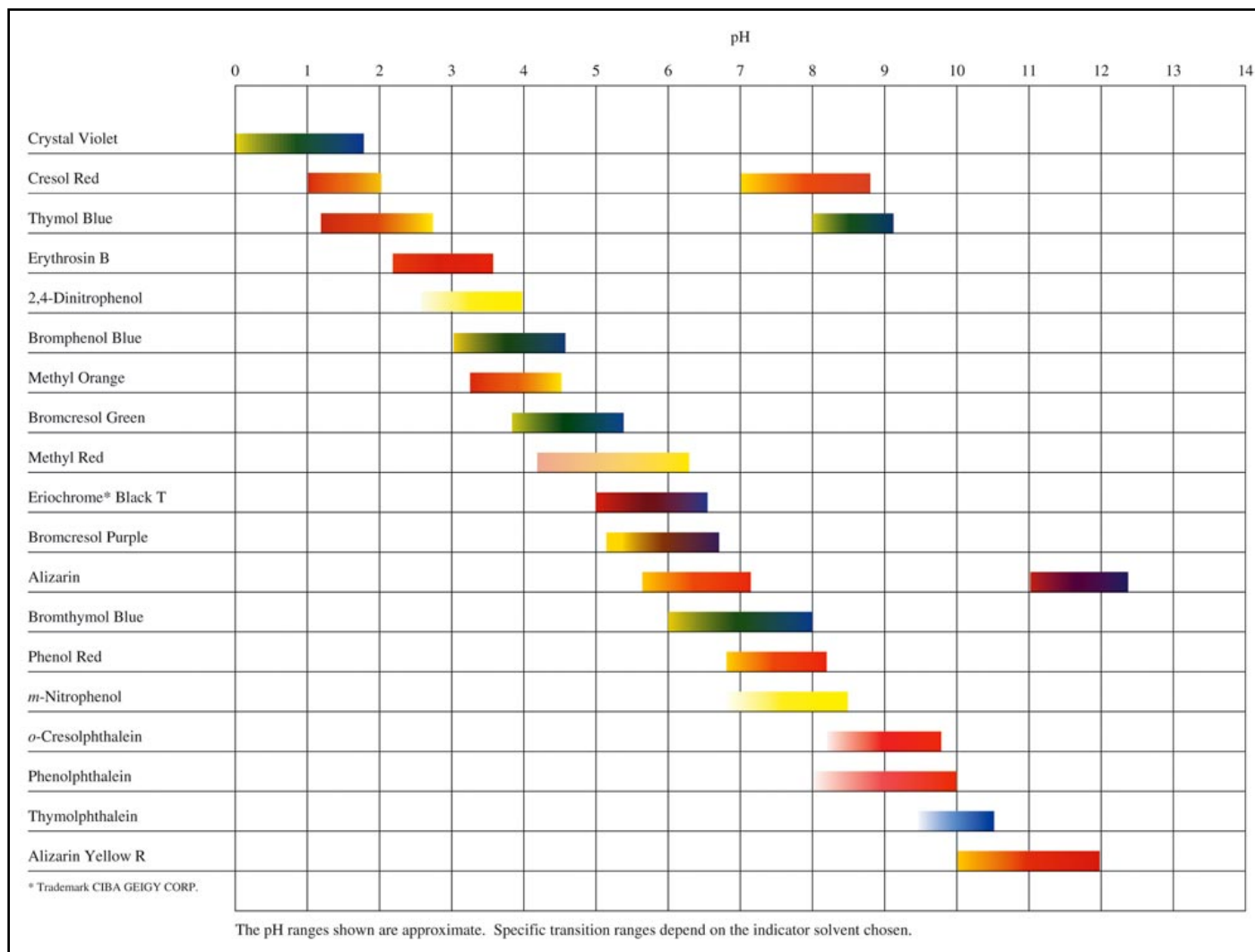


(Colorless acid form, HIn)



(Pink base form, In<sup>-</sup>)

Figure 15.8 The Useful pH Ranges for Several Common Indicators



Useful range:  $\text{pH} = \text{pK}_{\text{In}} \pm 1$  depending on Acid or Base Titration

Figure 15.9  
The pH Curve  
for the  
Titration of  
100.0 mL of  
0.10 *M* of HCl  
with 0.10 *M*  
NaOH

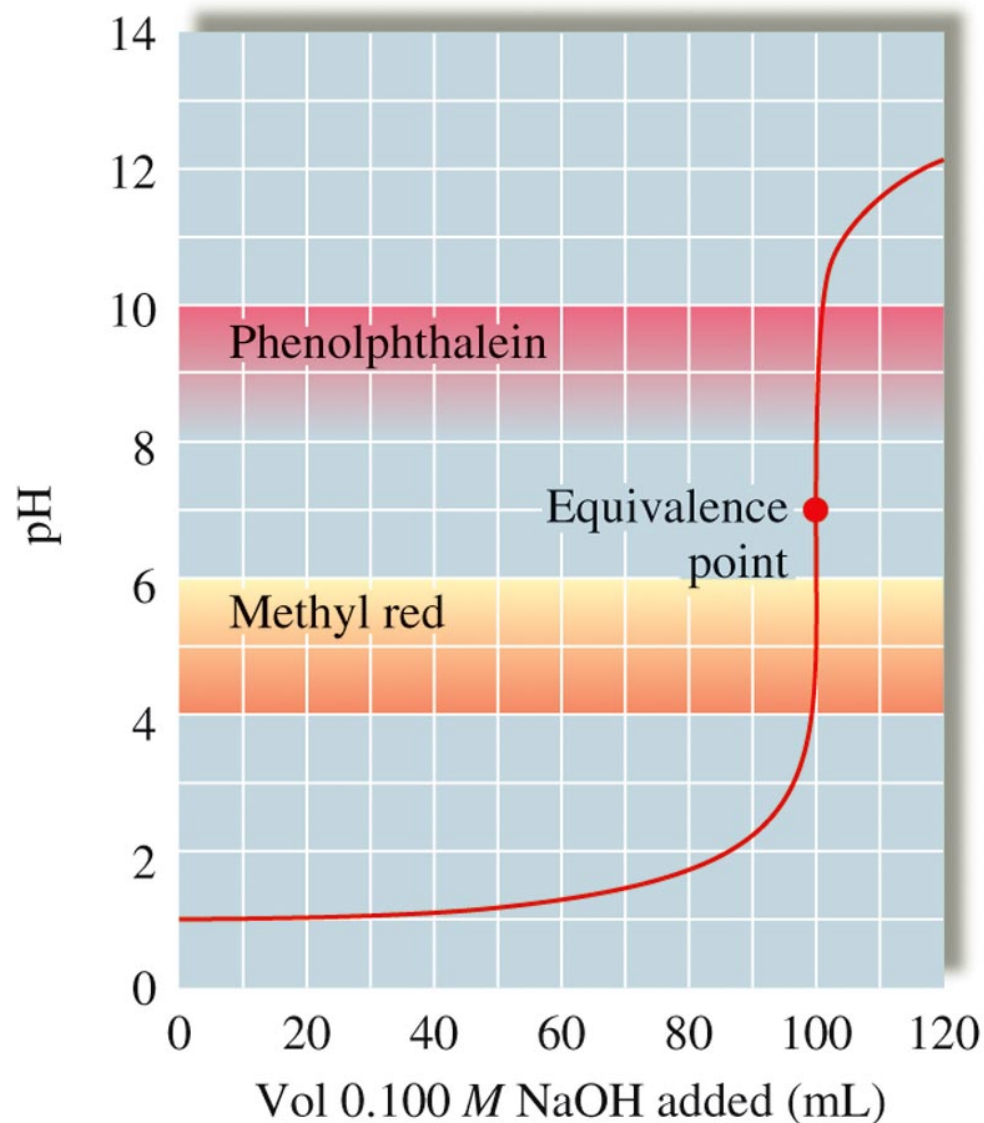
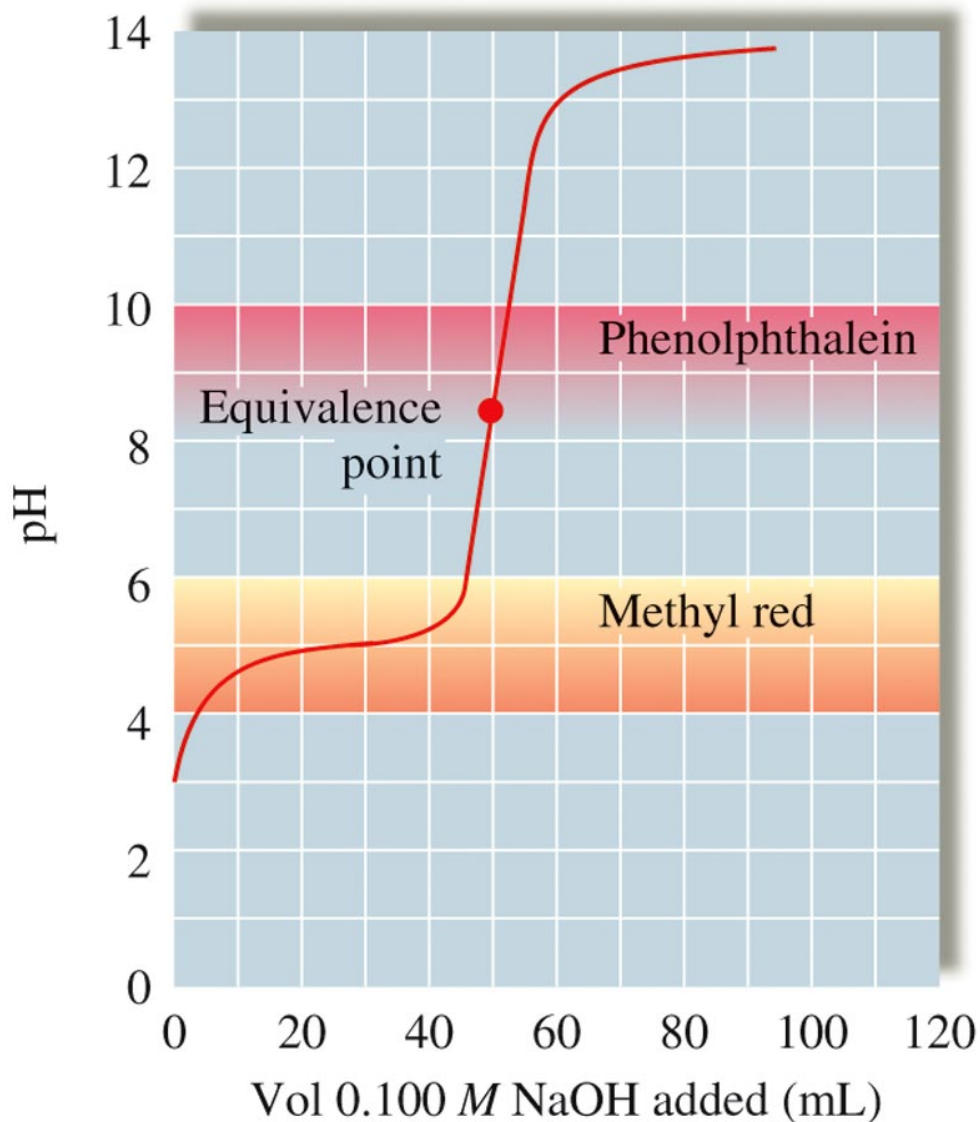


Figure 15.10  
The pH Curve for  
the Titration of 50  
mL of 0.1 *M*  
 $\text{HC}_2\text{H}_3\text{O}_2$  with 0.1  
*M* NaOH;  
Phenolphthalein  
Will Give an End  
Point Very Close  
to the  
Equivalence Point  
of the Titration



# Table 15.4 $K_{sp}$ Values at 25° Celsius for Common Ionic solids

**TABLE 15.4  $K_{sp}$  Values at 25°C for Common Ionic Solids**

Ionic Solid	$K_{sp}$ (at 25°C)	Ionic Solid	$K_{sp}$ (at 25°C)	Ionic Solid	$K_{sp}$ (at 25°C)
<b>Fluorides</b>		$\text{Hg}_2\text{CrO}_4^*$	$2 \times 10^{-9}$	$\text{Co}(\text{OH})_2$	$2.5 \times 10^{-16}$
$\text{BaF}_2$	$2.4 \times 10^{-5}$	$\text{BaCrO}_4$	$8.5 \times 10^{-11}$	$\text{Ni}(\text{OH})_2$	$1.6 \times 10^{-16}$
$\text{MgF}_2$	$6.4 \times 10^{-9}$	$\text{Ag}_2\text{CrO}_4$	$9.0 \times 10^{-12}$	$\text{Zn}(\text{OH})_2$	$4.5 \times 10^{-17}$
$\text{PbF}_2$	$4 \times 10^{-8}$	$\text{PbCrO}_4$	$2 \times 10^{-16}$	$\text{Cu}(\text{OH})_2$	$1.6 \times 10^{-19}$
$\text{SrF}_2$	$7.9 \times 10^{-10}$			$\text{Hg}(\text{OH})_2$	$3 \times 10^{-26}$
$\text{CaF}_2$	$4.0 \times 10^{-11}$	<b>Carbonates</b>		$\text{Sn}(\text{OH})_2$	$3 \times 10^{-27}$
		$\text{NiCO}_3$	$1.4 \times 10^{-7}$	$\text{Cr}(\text{OH})_3$	$6.7 \times 10^{-31}$
<b>Chlorides</b>		$\text{CaCO}_3$	$8.7 \times 10^{-9}$	$\text{Al}(\text{OH})_3$	$2 \times 10^{-32}$
$\text{PbCl}_2$	$1.6 \times 10^{-5}$	$\text{BaCO}_3$	$1.6 \times 10^{-9}$	$\text{Fe}(\text{OH})_3$	$4 \times 10^{-38}$
$\text{AgCl}$	$1.6 \times 10^{-10}$	$\text{SrCO}_3$	$7 \times 10^{-10}$	$\text{Co}(\text{OH})_3$	$2.5 \times 10^{-43}$
$\text{Hg}_2\text{Cl}_2^*$	$1.1 \times 10^{-18}$	$\text{CuCO}_3$	$2.5 \times 10^{-10}$		
		$\text{ZnCO}_3$	$2 \times 10^{-10}$	<b>Sulfides</b>	
<b>Bromides</b>		$\text{MnCO}_3$	$8.8 \times 10^{-11}$	$\text{MnS}$	$2.3 \times 10^{-13}$
$\text{PbBr}_2$	$4.6 \times 10^{-6}$	$\text{FeCO}_3$	$2.1 \times 10^{-11}$	$\text{FeS}$	$3.7 \times 10^{-19}$
$\text{AgBr}$	$5.0 \times 10^{-13}$	$\text{Ag}_2\text{CO}_3$	$8.1 \times 10^{-12}$	$\text{NiS}$	$3 \times 10^{-21}$
$\text{Hg}_2\text{Br}_2^*$	$1.3 \times 10^{-22}$	$\text{CdCO}_3$	$5.2 \times 10^{-12}$	$\text{CoS}$	$5 \times 10^{-22}$
		$\text{PbCO}_3$	$1.5 \times 10^{-15}$	$\text{ZnS}$	$2.5 \times 10^{-22}$
<b>Iodides</b>		$\text{MgCO}_3$	$6.8 \times 10^{-6}$	$\text{SnS}$	$1 \times 10^{-26}$
$\text{PbI}_2$	$1.4 \times 10^{-8}$	$\text{Hg}_2\text{CO}_3^*$	$9.0 \times 10^{-15}$	$\text{CdS}$	$1.0 \times 10^{-28}$
$\text{AgI}$	$1.5 \times 10^{-16}$			$\text{PbS}$	$7 \times 10^{-29}$
$\text{Hg}_2\text{I}_2^*$	$4.5 \times 10^{-29}$	<b>Hydroxides</b>		$\text{CuS}$	$8.5 \times 10^{-45}$
		$\text{Ba}(\text{OH})_2$	$5.0 \times 10^{-3}$	$\text{Ag}_2\text{S}$	$1.6 \times 10^{-49}$
<b>Sulfates</b>		$\text{Sr}(\text{OH})_2$	$3.2 \times 10^{-4}$	$\text{HgS}$	$1.6 \times 10^{-54}$
$\text{CaSO}_4$	$6.1 \times 10^{-5}$	$\text{Ca}(\text{OH})_2$	$1.3 \times 10^{-6}$		
$\text{Ag}_2\text{SO}_4$	$1.2 \times 10^{-5}$	$\text{AgOH}$	$2.0 \times 10^{-8}$	<b>Phosphates</b>	
$\text{SrSO}_4$	$3.2 \times 10^{-7}$	$\text{Mg}(\text{OH})_2$	$8.9 \times 10^{-12}$	$\text{Ag}_3\text{PO}_4$	$1.8 \times 10^{-18}$
$\text{PbSO}_4$	$1.3 \times 10^{-8}$	$\text{Mn}(\text{OH})_2$	$2 \times 10^{-13}$	$\text{Sr}_3(\text{PO}_4)_2$	$1 \times 10^{-31}$
$\text{BaSO}_4$	$1.5 \times 10^{-9}$	$\text{Cd}(\text{OH})_2$	$5.9 \times 10^{-15}$	$\text{Ca}_3(\text{PO}_4)_2$	$1.3 \times 10^{-32}$
		$\text{Pb}(\text{OH})_2$	$1.2 \times 10^{-15}$	$\text{Ba}_3(\text{PO}_4)_2$	$6 \times 10^{-39}$
<b>Chromates</b>		$\text{Fe}(\text{OH})_2$	$1.8 \times 10^{-15}$	$\text{Pb}_3(\text{PO}_4)_2$	$1 \times 10^{-54}$
$\text{SrCrO}_4$	$3.6 \times 10^{-5}$				

\*Contains  $\text{Hg}_2^{2+}$  ions.  $K = [\text{Hg}_2^{2+}][\text{X}^-]^2$  for  $\text{Hg}_2\text{X}_2$  salts, for example.

# QUESTION

Calcium phosphate ( $\text{Ca}_3(\text{PO}_4)_2$ ) is only slightly soluble. In fact, it is common ingredient in phosphate rock and is a major source of phosphate fertilizer. If the molar solubility is  $2 \times 10^{-7}$ , what is the value of the  $K_{\text{sp}}$ ?

$$\text{Ans} = 3 \times 10^{-32}$$

# QUESTION

Lead (II) iodide is used in some camera batteries.  $\text{PbI}_2$  has a  $K_{\text{sp}}$  of  $1.4 \times 10^{-8}$ . What is the molar solubility of this compound?

$$\text{Ans} = 1.5 \times 10^{-3} \text{ M}$$

## Table 15.5 Calculated Solubilities for CuS, Ag<sub>2</sub>S, and Bi<sub>2</sub>S<sub>3</sub> at 25° Celsius

**TABLE 15.5 Calculated Solubilities for CuS, Ag<sub>2</sub>S, and Bi<sub>2</sub>S<sub>3</sub> at 25°C**

Salt	$K_{sp}$	Calculated Solubility (mol/L)
CuS	$8.5 \times 10^{-45}$	$9.2 \times 10^{-23}$
Ag <sub>2</sub> S	$1.6 \times 10^{-49}$	$3.4 \times 10^{-17}$
Bi <sub>2</sub> S <sub>3</sub>	$1.1 \times 10^{-73}$	$1.0 \times 10^{-15}$

# QUESTION

The  $K_{\text{sp}}$  of  $\text{Ag}_2\text{CrO}_4$  is  $9.0 \times 10^{-12}$ . What would be the solubility of this compound in a solution that was already 0.10 M in potassium chromate?

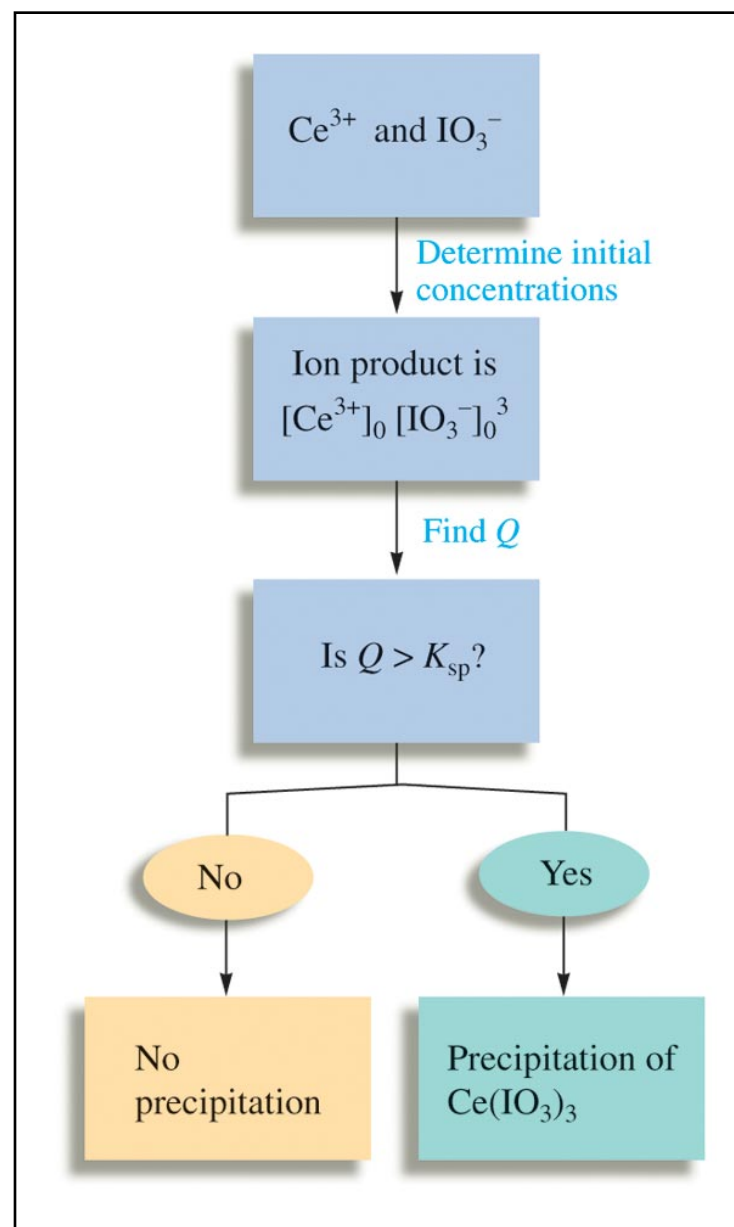
$$\text{Ans} = 4.7 \times 10^{-6} \text{ M}$$

# Determining Precipitation Conditions

A solution is prepared by adding 750.0 mL of  $4.00 \times 10^{-3} M$   $\text{Ce}(\text{NO}_3)_3$  to 300.0 mL of  $2.00 \times 10^{-2} M$   $\text{KIO}_3$ . Will  $\text{Ce}(\text{IO}_3)_3$  ( $K_{\text{sp}} = 1.9 \times 10^{-10}$ ) precipitate?

You Betcha!

What are the [ ]'s of each ion after precipitation?



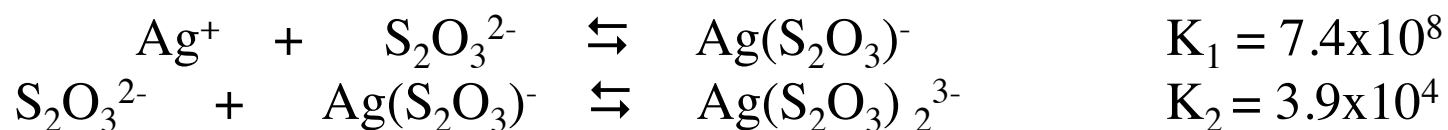
# QUESTION

To separate a solution containing 0.00010 M silver and 0.10 M lead ions, as may be done in qualitative analysis separation schemes, a source of  $\text{I}^-$  may be slowly added to the mixture of ions. Which will precipitate first:  $\text{AgI}$  ( $K_{\text{sp}} = 1.5 \times 10^{-16}$ ) or  $\text{PbI}_2$  ( $K_{\text{sp}} = 1.4 \times 10^{-8}$ )? Also, what would be the concentration of  $\text{I}^-$  necessary to see that first precipitation?

Ans =  $\text{AgI}$ ;  $[\text{I}^-]$  would be  $1.5 \times 10^{-12} \text{ M}$

# QUESTION

Calculate the concentrations of  $\text{Ag}^+$ ,  $\text{Ag}(\text{S}_2\text{O}_3)^-$ , and  $\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}$  in a solution prepared by mixing 150.0 mL of  $1.00 \times 10^{-3} \text{ M}$   $\text{AgNO}_3$  with 200.0 mL of  $5.00 \text{ M}$   $\text{Na}_2\text{S}_2\text{O}_3$ . The stepwise formation equilibria are:



$$[\text{Ag}(\text{S}_2\text{O}_3)_2^{3-}] = 4.29 \times 10^{-4} \text{ M}$$

$$[\text{Ag}(\text{S}_2\text{O}_3)^-] = 3.8 \times 10^{-9} \text{ M}$$

$$[\text{Ag}^+] = 1.8 \times 10^{-18} \text{ M}$$